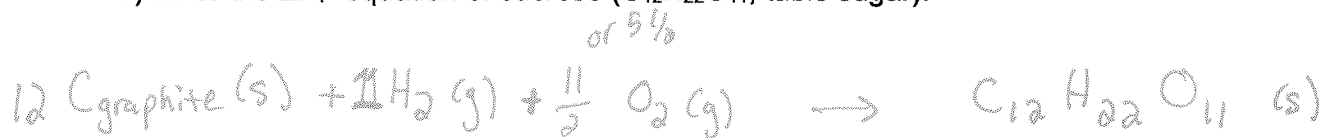
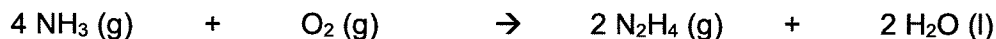


Chapter 5 – Extra Practice with Hess's Law and Enthalpy

1) Write the ΔH_f° equation of sucrose ($C_{12}H_{22}O_{11}$, table sugar).



2) Calculate the standard enthalpy change for the following reaction:



Chemical Species	ΔH_f° (kJ/mol)
H ₂ O (l)	-285.8
H ₂ O (g)	-241.8
O (g)	247.5
NH ₃ (g)	-46.19
NH ₃ (aq)	-80.29
N ₂ H ₄ (g)	95.40

$\Delta H_{\text{rxn}}^\circ = \text{pds} - \text{reactants}$

$\Delta H_{\text{rxn}}^\circ = [(2 \text{ N}_2\text{H}_4(\text{g})) + (2 \text{ H}_2\text{O}(\text{l}))] - [4(\text{NH}_3(\text{g})) + \text{O}_2(\text{g})]$

$\Delta H_{\text{rxn}}^\circ = \left[2 \left(\frac{95.40 \text{ kJ}}{\text{mol}} \right) + 2 \left(\frac{-285.8 \text{ kJ}}{\text{mol}} \right) \right] - \left[4 \left(\frac{-46.19 \text{ kJ}}{\text{mol}} \right) - 0 \right]$

element in standard state

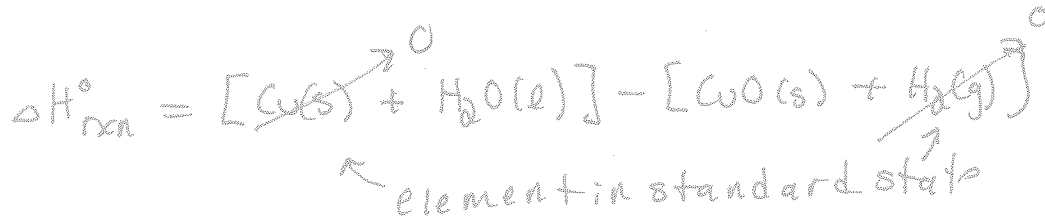
$\Delta H_{\text{rxn}}^\circ = [190.8 \text{ kJ} + -571.6 \text{ kJ}] - [-184.8 \text{ kJ}]$

$= -380.8 \text{ kJ} + 184.8 \text{ kJ}$

$\Delta H_{\text{rxn}}^\circ = -196.04 \text{ kJ} = \boxed{-196.0 \text{ kJ}}$

Given: $\text{CuO (s)} + \text{H}_2 \text{(g)} \rightarrow \text{Cu (s)} + \text{H}_2\text{O (l)} \quad \Delta H_{\text{rxn}}^\circ = -129.7 \text{ kJ}$

Chemical Species	ΔH_f° (kJ/mol)
H ₂ O (l)	-285.8
H ₂ O (g)	-241.8
O (g)	247.5



$$\Delta H_{rxn}^{\circ} = [H_2O(l)] - [CO(g)]$$

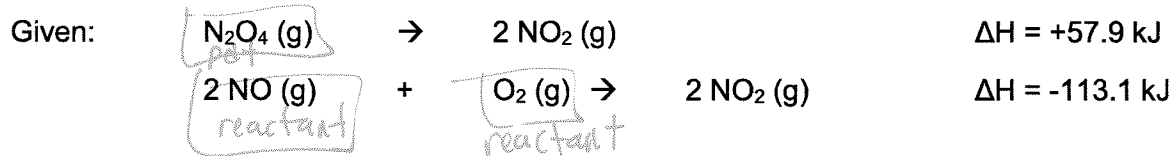
$$[\Delta H_f^\circ [\text{H}_2\text{O}(\ell)]] - \Delta H_{\text{rxn}}^\circ = [\Delta H_f^\circ [\text{CuO}(\text{s})]]$$

$$\Delta H_f^\circ \text{CO}(s) = +120.2 \text{ kJ} [-285.8 \text{ kJ}] \leftarrow [-129.7 \text{ kJ}]$$

$$\Delta H_f^\circ \text{CuO(s)} = -156.1 \text{ kJ/mol}$$

← b/c 2 mol from balanced chemical eq.

4) Calculate ΔH for: $2 \text{ NO (g)} + \text{O}_2 \text{ (g)} \rightarrow \text{N}_2\text{O}_4 \text{ (g)}$
reactant *reactant* *product*



$$\Delta H = \text{Vaporizing substance} (-57.9 \text{ kJ}) + (-113.1 \text{ kJ})$$

$$\Delta H = 171.0 \text{ KJ}$$